Chemistry 129.01 - General Chemistry Workshop - Spring 2017

Week #4

Monday, February 13. Sessions 4 of the Greenhouse Gas Module.
Assigned reading: Sections 7.1 – 7.4

The session 4 question is: How are the electrons distributed in greenhouse gases? We will spend some of today and a lot of Wednesday practicing drawing Lewis structures for covalent and ionic compounds and trying to understand the difference between them.

Sections 7.1 and 7.2 are critical for our understanding of what a "bond" is. After reading them, define in your own words what a chemical bond is. Write down some bonds that you would expect to have a covalent bond or an ionic bond. Section 7.2 also talks about electronegativity. Look at the periodic table in Figure 7.6. Look at the trend of electronegativities as you move from Li to F on the first row; which end is higher? Look at the trend as you go down the halogen group (F, Cl, Br, I, At); which end is higher? Based on these two trends, where should the least electronegative element be? Find it.

Section 7.4 introduces Lewis structures as a simple way to represent the valence shell of electrons and chemical bonds. Page 310 goes over the rules you need to know to draw Lewis Structures of molecules. Try your hand at a few of the examples before coming to class and generate some questions you can ask to understand this better.

Before Monday's class,

1. Define the following terms:

   Covalent Bond

   Ionic Bond

   Electronegativity

   Octet rule

2. Draw Lewis structure of:
   a. F atom. How many valence electrons does it have? How many does it need to have a full octet?
   b. F_2 molecule. How many electrons are shared? Is this a single, double or triple bond?

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Friday, February 17. Finish Session 8 of GHG Module

Assigned reading: Section 7.6

Quiz today: Orbitals, quantum numbers, electron configurations, periodic properties.

Lewis structures a simple way to represent molecules on paper but it doesn't provide any information about the 3D arrangement of groups around the central atom. We'll continue to practice drawing Lewis structures of molecules and predict their shape using the VSEPR model. Table 7.16 and 7.19 list the electron-group and molecular geometries, respectively. We'll discuss how to assign these shapes based on the Lewis structures of the molecules as well as how to determine if a molecule is polar or nonpolar based on its shape.

Before Friday's class,

1. Define the following terms:

   Bonding Electron-Group

   Nonbonding Electron-Group

   Dipole Moment

2. Name the electron-group geometries listed on table 7.16. What happens to the bond angles as the number of electron-groups increase?

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Problem Set #1

Due Monday, February 20 (at the beginning of class). Late homework will not be accepted.

1. Write a Lewis structure for each atom or ion:
   a. Al
   b. Mg^{2+}
   c. Cl
   d. Cl^{-}

2. Determine whether bond between each pair of atoms would be nonpolar, polar or ionic. Which atom is the more electronegative of the pair?
   a. Br and Br
   b. C and Cl
   c. Sr and O
   d. H and F

3. Write a Lewis structure for each ionic compound:
   a. NaF
   b. CaO
   c. $SrBr_2$
   d. K_2O

4. Write a Lewis structure for each molecule or ion. Which structures do not follow the octet rule?
   a. PH_3
   b. BCl_3
   c. HI
   d. NH_4^+
   e. SF_4
   f. SiH_4
   g. XeF_4
   h. CO_3^{2-}

5. The cyanate ion (OCN^-) has three possible Lewis structures. Draw the structures and determine which structure is likely to contribute most to the correct structure of OCN^-?

6. The three species NH_2^-, NH_3 and NH_4^+ have H—N—H bond angles of 105°, 107°, and 109.5°, respectively. Explain this variation in bond angles.

7. An AB_4 molecule is described as having a square planar geometry. How many nonbonding groups are on atom A? Explain.

8. Draw the Lewis structures of the following molecules or ions and give their <u>electron-group geometry</u> and <u>molecular geometry</u>:
   a. HCN
   b. SF_4
   c. NO_2^-
   d. NI_3
   e. SO_3
   f. XeF_2
   g. SF_6
   h. ClO_4^-